1.2 Chemistry of Acids and Bases

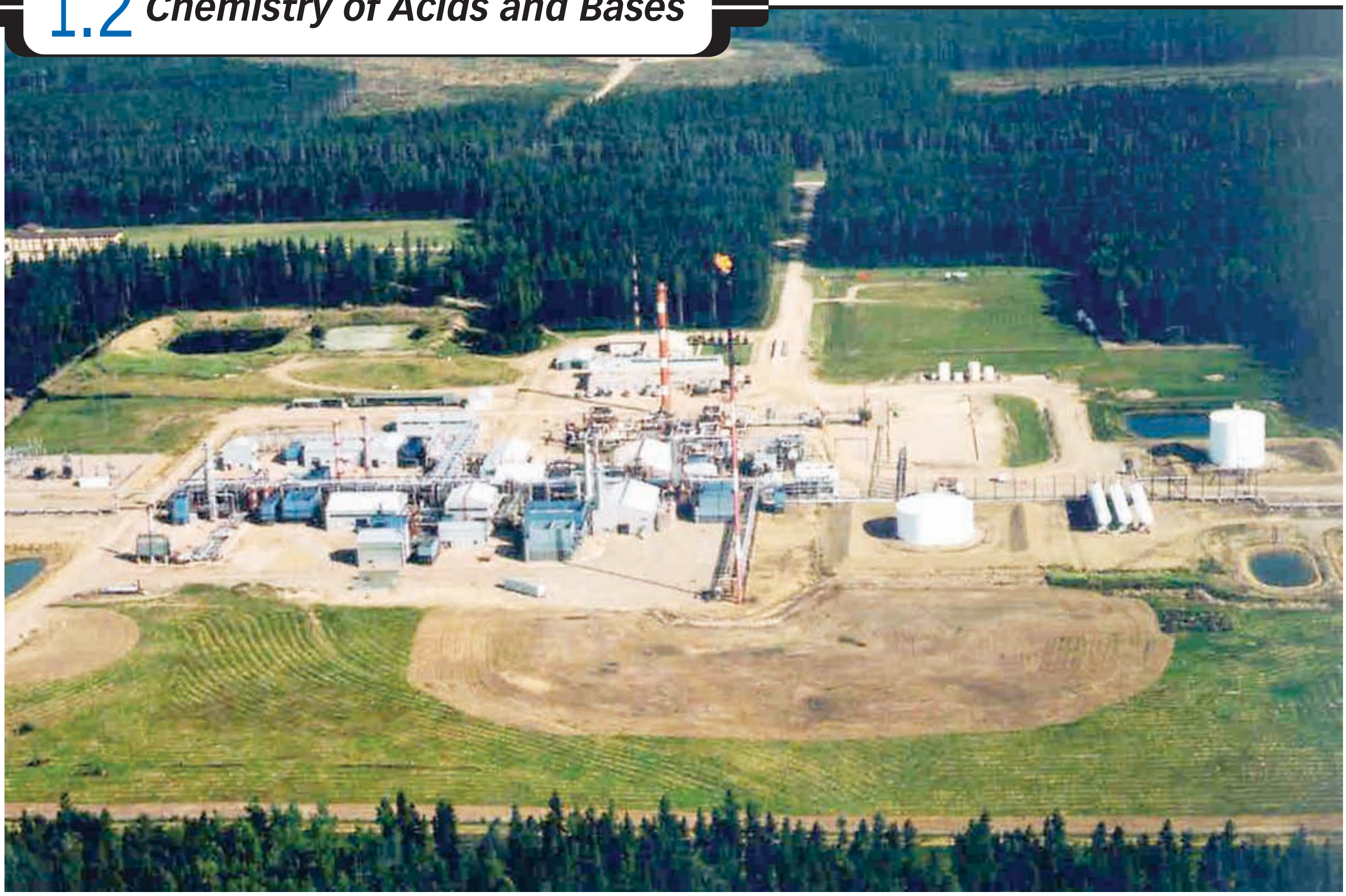


Figure B1.10

Most of the natural gas collected and processed from gas wells in Alberta contains hydrogen sulfide, $H_2S(g)$. At sour gas processing facilities (like the one shown in Figure B1.10), the hydrogen sulfide is removed and is converted into sulfur. Metal pipes, like the one shown in Figure B1.11, can be damaged by exposure to sour gas. Recall from previous science courses that the corrosion of metal objects occurs when certain substances come into contact with one another.

Workers in the oil and gas industry continually monitor the corrosion of pipes, especially where sour gas is extracted and processed. Strict safety standards are maintained to protect people and the environment from the effects of sour gas.

In this lesson you will closely study the chemical components within sour gas that enable it to affect metals and other substances. You will also see how the products of combustion reactions can further react to produce acidic solutions. You will examine how acidic solutions can affect other substances, including bases, within chemical systems. You will then discover methods to measure and describe the acidity of solutions.

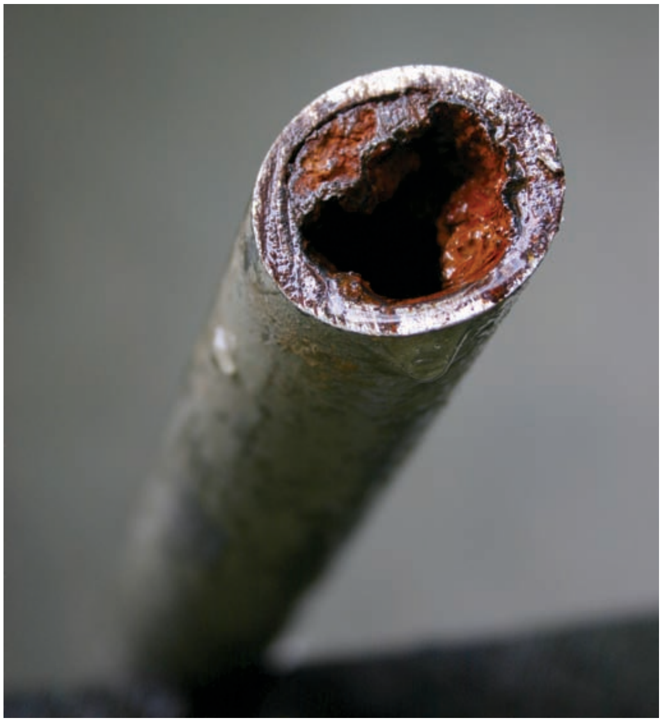


Figure B1.11: A metal pipe damaged by corrosion

What Makes Sour Gas Sour?

Figure B1.12: The sour taste of many foods is distinctive and is produced by the acids naturally present within the food or by acids added during its production.

Words like *sour* are used to provide a description. Oranges, lemon juice, or perhaps your favourite candy can be described as having a sour taste. Likewise, scientists use descriptive terms to communicate the behaviour of substances they investigate. Often, these descriptions are based on observations of the matter during experimentation. Descriptions of the response of substances to tests performed during an experiment can be used to make **empirical** definitions.

As described earlier, sour gas contains hydrogen sulfide, $H_2S(g)$. Since water is often present within pipes containing sour gas, hydrogen sulfide can dissolve into water to form an **aqueous solution**, represented as $H_2S(aq)$.

empirical: a result of an observation

aqueous solution: a solution in which water is the solvent

Practice

9. In previous science courses you were introduced to the terms *ionic compound*, *molecular compound*, *acid*, and *base*. Match each term with one of the following definitions.
   a. a compound composed of oppositely charged particles, often metal and non-metal atoms; usually forms conductive solutions, called electrolytes, when dissolved in water
   b. a corrosive solution containing hydrogen in the chemical formula of the solute
   c. a compound composed of two or more non-metal atoms; may dissolve in water; only some form electrolytes, but most often form non-electrolytes
   d. a caustic, corrosive solution; often contains a hydroxide ion and a group 1 or 2 metal in the periodic table
10. a. In previous science courses you worked with conductivity meters and litmus paper (red and blue litmus). Familiarize yourself with how and why each of these pieces of equipment is used. Summarize your review by copying and completing the following table.

Apparatus	Used to Identify	Expected Result for a Positive Test	Expected Result for a Negative Test
conductivity meter			
red litmus paper			
blue litmus paper			

   b. Describe the experimental control used with each apparatus. Include the solution tested and the expected result for the test.

Empirical Properties of Acids, Bases, and Neutral Solutions

Careful planning and attention to detail is essential when you design an experiment or investigation. Your attention to detail, such as performing the same test on experimental controls, improves the quality of the data you collect. The quality of data is determined by considering validity and reliability when designing and performing the investigation. The “Reliability and Validity” table on page 167 demonstrates some questions and actions you might take to make your experimental design and procedure reliable and valid. By paying attention to these aspects, you will become more confident in your data.

RELIABILITY AND VALIDITY

Reliability	Validity
Questions the Way Experiment Is Performed Is it possible to obtain the same result if I repeat the experiment using the same method?	**Questions the Process Used to Obtain Measurements** Does this process measure what it is supposed to?
How to Improve Reliability • Repeat tests with both positive and negative experimental controls. • Perform frequent calibration checks. • Practise techniques and use of equipment.	**How to Improve Validity** • Select equipment that is appropriate for the experiment. • Select methods that others have used successfully to perform similar tasks. • Use the equipment appropriately.

In the next investigation you will test some aqueous solutions and describe the response of each solution to the tests. Once you are finished, you should be able to identify trends within your observations and develop a set of empirical definitions for the solutions tested.

Investigation

Testing Aqueous Solutions

Science Skills
- ✓ Initiating and Planning
- ✓ Performing and Recording

Purpose

You will design and perform an experiment to identify acidic, basic, neutral molecular, and neutral ionic solutions.

Materials

- 0.100-mol/L solutions of
  - $HCl(aq)$
  - $HNO_3(aq)$
  - $H_2SO_4(aq)$
  - $H_2S(aq)$
  - $NaOH(aq)$
  - $Na_2CO_3(aq)$
  - $Na_2SO_4(aq)$
  - $NaCl(aq)$
  - $CH_3OH(aq)$
- distilled or de-ionized water
- multiwell dish (or 13 watch glasses or Petri dishes)
- blue and red litmus paper
- magnesium turnings (or iron filings)
- stirring rod
- forceps or tweezers
- MSDS (Material Safety Data Sheet) for each solution
- conductivity meter (or tester)

Procedure

step 1: Develop an experimental design that includes the following considerations:

- **Safety:** Identify solutions that contain compounds that are irritants or that may cause some other safety concern. Consult the MSDS information for each solution.
- **Manipulation of apparatus:** If necessary, seek further instruction from your teacher regarding the use of the apparatus (e.g., the conductivity meter).
- **Cleanup:** Learn the proper procedure for the disposal of the chemicals. Determine how the apparatus should be cleaned.

step 2: Have your teacher approve your procedure before you begin.

step 3: Follow your procedure, and record your results.

CAUTION!

Use gloves, safety glasses, and a lab apron for this activity.

Analysis

1. Identify the positive and negative controls in the investigation.
2. Identify actions taken during the investigation that improved the quality of the data collected.
3. Describe how the data collected during the investigation demonstrates reliability.
4. Describe how the tests completed during the investigation address validity.

Trends and Patterns in Data

Trends and patterns within experimental data are important. When you look at the data from the "Testing Aqueous Solutions" investigation, do you notice that some of the solutions behaved in a similar manner? Is it possible to sort the solutions using the similarities in their behaviour to certain tests?

Did you notice how many of the solutions had a similar reaction to each type of litmus paper used in the tests? The sorting of substances based on their similar behaviours to certain tests was used to create the empirical definitions you explored earlier. Refer to the "Properties of Acids, Bases, and Neutral Solutions" table.

PROPERTIES OF ACIDS, BASES, AND NEUTRAL SOLUTIONS

Solution	Properties
acid	• electrolytic (conducts a current) • corrosive • turns blue litmus red • reacts with active metals (e.g., Mg, Zn, and Fe) to produce hydrogen gas • neutralized by bases and basic solutions • tastes sour
base	• electrolytic (conducts a current) • corrosive • turns red litmus blue • feels slippery (when diluted) • neutralized by acids and acidic solutions • tastes bitter
neutral	• can be electrolytic (if solute is an ionic compound) • does not change red or blue litmus

Types of Deposition

Emissions from industrial activities can carry sulfur dioxide and other substances great distances. A major factor in determining how long emissions will stay in the atmosphere is how quickly they come into contact with other materials in the environment. Emissions that contact liquid or solid forms of water in the atmosphere can dissolve and return as **wet deposition**. Gases and particles within emissions that are absorbed by Earth's surface are called **dry deposition**. Alberta has a dry climate. It is estimated that most of the pollution from emissions in Alberta occurs in the form of dry deposition. The terms *wet* and *dry* refer to the state of the material being deposited; therefore, it is possible for dry deposition to be deposited onto any surface, including bodies of water (e.g., lakes and rivers).

wet deposition: gases or particles that are removed from the atmosphere by water (liquid or solid) and deposited as precipitation

dry deposition: gases or particles that are transported by winds and absorbed by Earth's surface

Practice

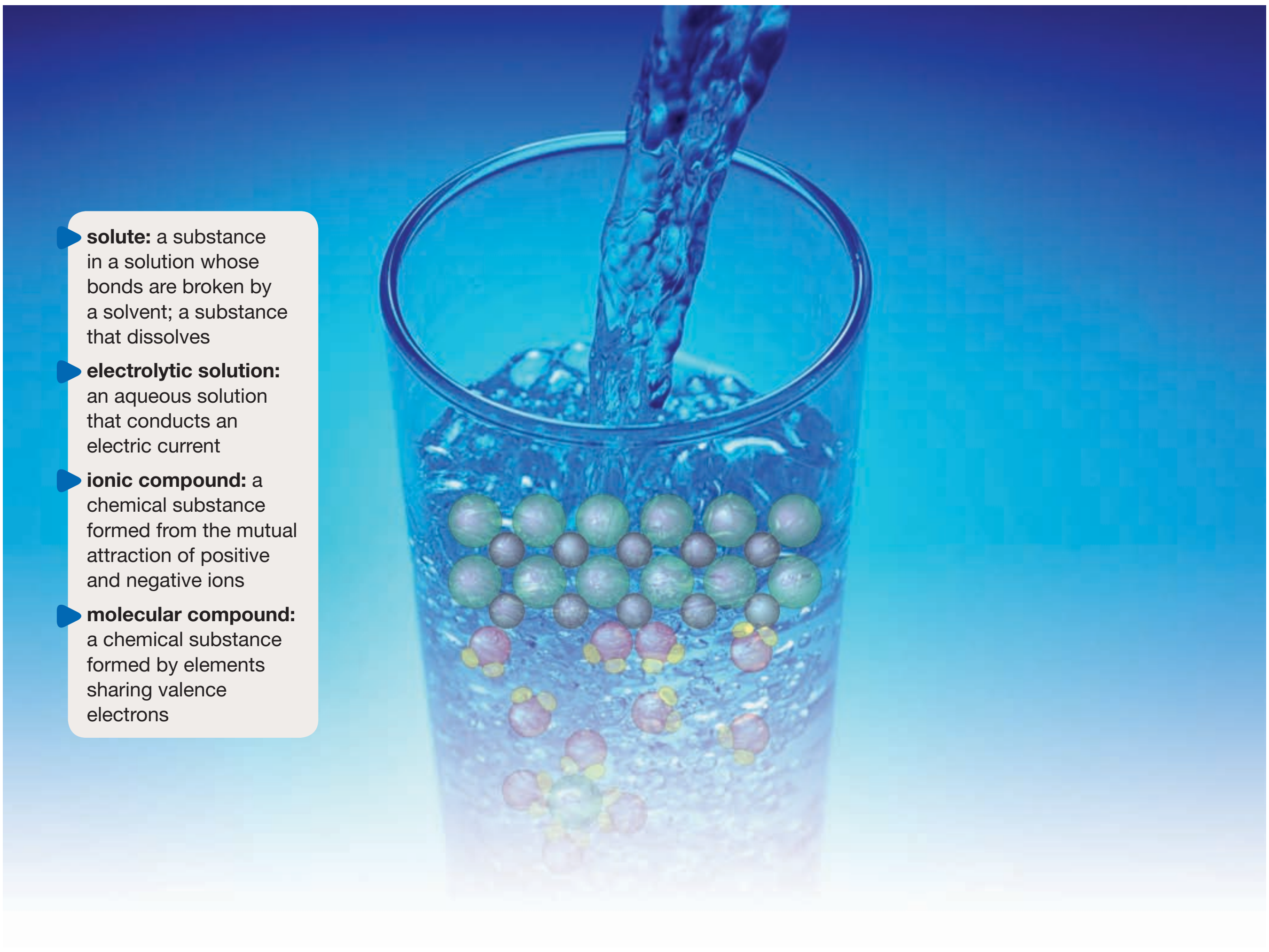
11. Copy and complete the following table to summarize the results from the "Testing Aqueous Solutions" investigation. For now, do not complete the Definition column.

Solution	Definition	Empirical Properties	Examples
acidic			
basic			
neutral			

What Makes a Solution Acidic?

Earlier, you were able to use similarities within your observations to classify a solution as being acidic, basic, or neutral. You also saw that the groupings you made coincided with the known empirical properties for acids, bases, and neutral solutions. Apart from these similarities, did you note any other similarities among the acidic solutions?

Acids are a special group of chemical compounds. Did you notice that all of the substances categorized as acids contain hydrogen and were dissolved in water? You may have also noted that all the acidic solutions tested were **electrolytic solutions**, even if their chemical formula suggests that the **solute** is a **molecular compound**. Although many acids are molecular compounds, all acids appear to behave like **ionic compounds** when dissolved in water. Acids tend to form electrolytic solutions, whereas molecular compounds form non-electrolytic solutions. As you will soon see, the microscopic changes that occur within a solution containing a dissolved acid are important when explaining the properties of acids.



solute: a substance in a solution whose bonds are broken by a solvent; a substance that dissolves

electrolytic solution: an aqueous solution that conducts an electric current

ionic compound: a chemical substance formed from the mutual attraction of positive and negative ions

molecular compound: a chemical substance formed by elements sharing valence electrons

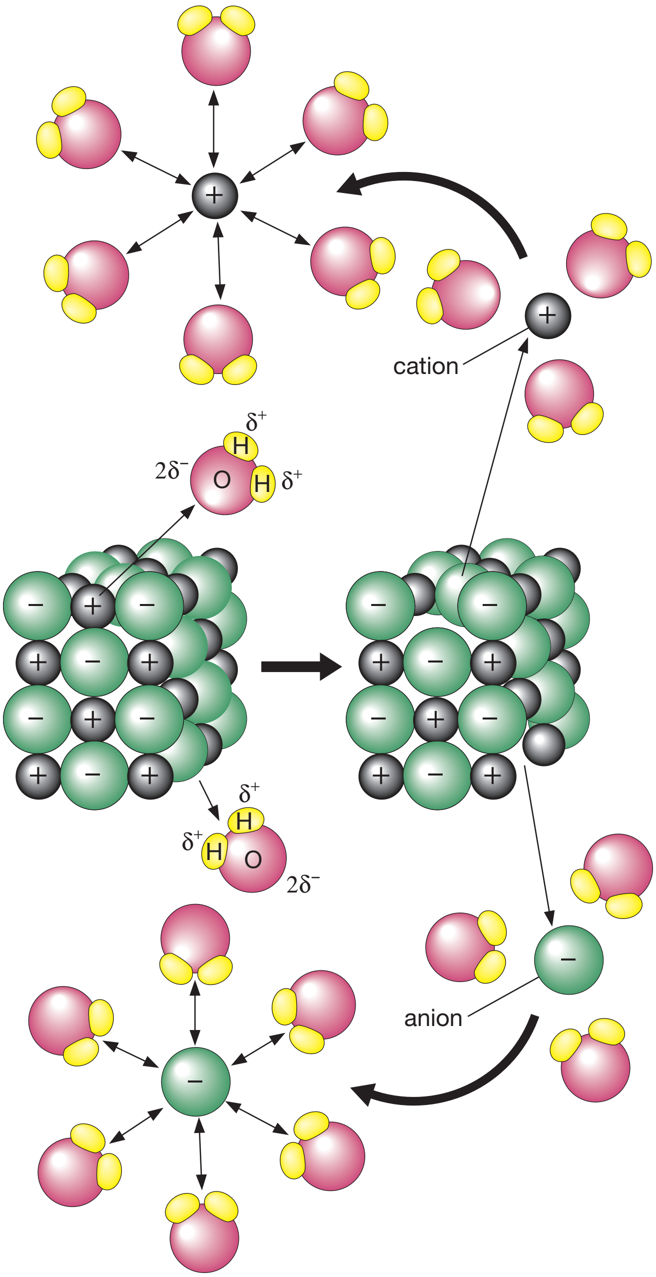
Acids and Bases in Solutions

Conductive solutions contain freely moving ions. In previous science courses you discovered that water molecules break the bonds between ions in an ionic solute, causing the ions to dissociate. **Dissociation** occurs due to **electrostatic attraction** between the charged ions of the solute and the charges on water molecules. Although dissociation cannot be seen, a positive conductivity test—implying that charged particles are present and are able to move within the solution being tested—is indirect evidence of this microscopic change. In 1834, an English physicist by the name of Michael Faraday was the first scientist to demonstrate that acids, bases, and salts (later determined to be composed of ionic compounds) all dissolve in water to form electrolytes.

dissociation: the separation of a chemical substance into its individual ions in a solution

electrostatic attraction: a force that acts to pull oppositely charged objects toward each other

Water Molecules Dissolving an Ionic Crystal



Science Links

Many phenomena, including lightning, result from electrostatic attraction between oppositely charged particles. In Unit C you will study the fields that surround objects and how forces like electrostatics are the product of fields.

Conductivity tests have demonstrated that all acidic solutions also conduct an electric current, indicating that the acid molecule has formed ions. Is this observation connected to the fact that all acids contain hydrogen? Some scientific theories attempting to explain the properties and behaviours of acids have focused on the ability of acids to form hydrogen ions in water.

In 1887, a Swedish chemist named Svante Arrhenius published a theory that suggested that acids form aqueous solutions that contain hydrogen ions, $H^+(aq)$, and a negatively charged ion. His theory also proposed that bases form solutions that contain hydroxide ions, $OH^-(aq)$, and a positively charged ion. Although not defined by this theory, solutions that produced neither a hydrogen ion nor a hydroxide ion can be considered neutral electrolytic solutions.

Figure B1.13: Svante Arrhenius (1859–1927)

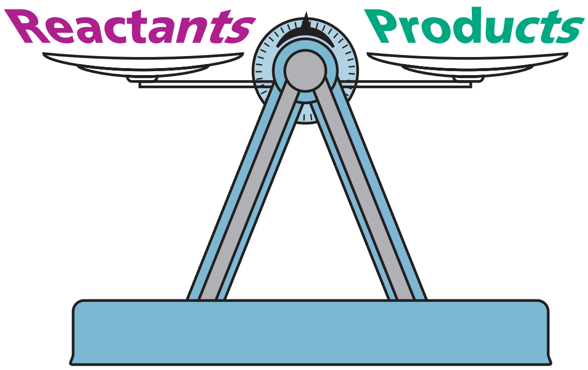
Changes to Solutes in Aqueous Solutions

Acids: e.g., hydrochloric acid
$HCl(aq) \rightarrow H^+(aq) + Cl^-(aq)$

Bases: e.g., potassium hydroxide
$KOH(aq) \rightarrow K^+(aq) + OH^-(aq)$

Neutral Substances: e.g., calcium chloride
$CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2\ Cl^-(aq)$

Balancing Chemical Equations



Balancing a chemical equation using coefficients demonstrates that all the atoms on the reactants side of the equation have been accounted for on the products side. Recall that matter cannot be created nor destroyed. Coefficients represent the number of each particle involved. When properly balanced, the net charge on each side of the equation will be the same.

Balancing Equations

$HCl(aq) \rightarrow H^{+}(aq) + Cl^{-}(aq)$

H = 1, Cl = 1
net charge = 0

H = 1, Cl = 1
net charge = 1(1+) + 1(1–)
= 0

coefficient needed to balance chloride ions

$CaCl_2(aq) \rightarrow Ca^{2+}(aq) + 2\ Cl^{-}(aq)$

Ca = 1, Cl = 2
net charge = 0

Ca = 1, Cl = 2
net charge = 1(2+) + 2(1–)
= 0

Chemical equations are balanced when

- an equal number of each type of atom appears on each side of the equation
- the net charge on each side of the equation is equal

Practice

12. Write a balanced equation for the change that occurred with each substance when it was dissolved in water.

a. $HNO_3(aq)$
b. $H_2SO_4(aq)$
c. $H_2S(aq)$
d. $NaOH(aq)$
e. $Na_2CO_3(aq)$
f. $Na_2SO_4(aq)$
g. $NaCl(aq)$

13. Use the equations written in question 12 to predict whether each solution listed is acidic, basic, or neutral. List any inconsistencies.

Limitations to Arrhenius's Theory

As previously stated, when chemical substances separate into their individual ions in a solution, this is called dissociation. Dissociation equations like that for $Na_2CO_3(aq)$, may be written to explain how ionic solutions can conduct an electric current.

$$Na_2CO_3(aq) \rightarrow 2Na^{+}(aq) + CO_3^{2-}(aq)$$

Arrhenius's theory states that the presence of hydroxide ions in the chemical formula of a solute explain its basic properties. $Na_2CO_3(aq)$, in solution, has definite basic properties. However, you can see that the above equation does not show the presence of $OH^{-}(aq)$ ions. How can these ions be responsible for basic properties? A similar problem exists for substances like $AlCl_3(aq)$, which is acidic in solution.

A second problem with Arrhenius's theory focuses on the possible existence of a free hydrogen ion moving among water molecules in a solution. Scientists questioned the possibility of free hydrogen ions existing within an aqueous solution. The simplest element is hydrogen, composed of one proton and one electron. Hydrogen atoms become positively charged when the single electron is removed. The absence of any electron, combined with the small size of hydrogen's atomic nucleus, results in the hydrogen ion having a very strong positive charge.

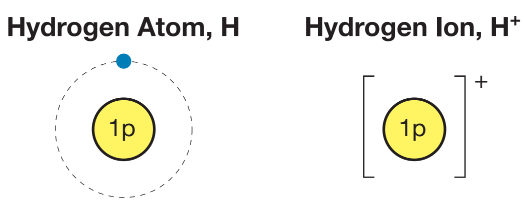


Figure B1.14: *Hydrogen ion*, *proton* . . . there are times when scientific terms can seem confusing. Can you explain why scientists sometimes refer to a hydrogen ion as a proton?

Many of the unique properties of a water molecule, including its ability to dissociate a solute, are explained by the **polarity** of the water molecule. Part of the polarity of the water molecule is due to exposed pairs of electrons located on the molecule's surface. Since electrons have a negative charge, the areas on the surface of a water molecule where these pairs of electrons are located will also have a partial negative charge. An electrostatic attraction between the positively charged hydrogen atom and the negatively charged areas of the water molecule provides an opportunity for these two objects to combine. The hydrogen ion becomes bound to the water molecule. The product of this reaction is the **hydronium ion**, $H_3O^+(aq)$.

polarity: the presence of different regions of charge on a molecule

hydronium ion: an ion created when a water molecule combines with a hydrogen ion; $H_3O^+(aq)$

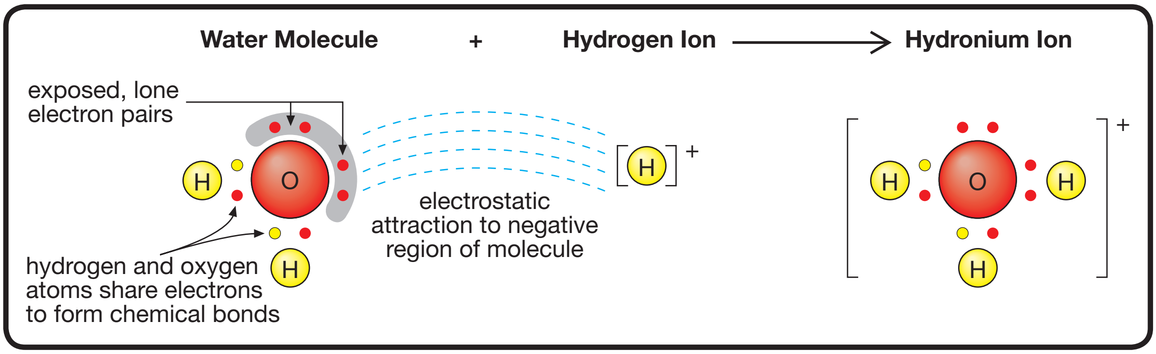


Figure B1.15: The creation of hydronium

Since the development of Arrhenius's theory, evidence has supported the existence of the hydronium ion, $H_3O^+(aq)$, within aqueous solutions. Currently, the hydronium ion is recognized as the acidic particle.

Apart from clarifying that the hydronium ion is the particle responsible for the acidity of solutions, later theories emphasized the importance of collisions between substances, including water, within a chemical system. In previous science courses, water was referred to as an excellent medium for chemical change that enables other particles to collide. Now, when you think of reactions within a solution, you may find that water is one of the reactants.

Exchange of Hydrogen Ions

You may have wondered if a collision between a hydronium ion and a water molecule could result in the transfer of the hydrogen ion to other water molecules or to other substances. The possibility of transferring hydrogen ions between substances within a solution stimulated the development of other theories to explain the behaviour of acids and bases.

In 1923, a Danish chemist—Johannes Brønsted—and an English chemist—Thomas Lowry—independently published similar theories that described an alternate way of explaining the behaviour of acids and bases. The coincidence of two scientists working in the same area of research and publishing similar theories was rare at the time. If you think this coincidence is strange, at the time Brønsted and Lowry did their work, there was no e-mail, Internet, or jet-airplane travel. Back then, the opportunities for scientists to meet and exchange ideas were greatly limited. The significance of these two researchers independently developing the same theory was great.

Scientific ideas and discoveries undergo a process of peer review that is designed to ensure the reliability of the experimentation, the validity of the data, and its interpretation. Peer review is one way scientists can check the work of others. During peer review, scientists analyze the details of how the data was collected, the data itself, the methods used to interpret data, and the ideas and theories developed from the interpretation of the data. Often, during the peer-review process, the reviewers may suggest that further experimentation is needed to provide better support for the conclusions.

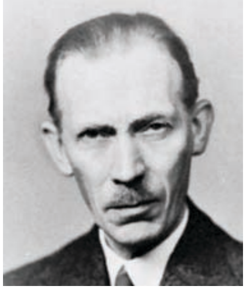


Figure B1.16: Johannes Brønsted (1879–1947)


Figure B1.17: Thomas Lowry (1874–1936)

DID YOU KNOW?

In their haste to publish a scientific discovery of cold fusion, Stanley Pons and Martin Fleishman chose to share their experimental findings with the media before the peer review. They made headlines twice: once for the discovery of cold fusion—a process believed to produce energy—and the other after their work was peer reviewed and found to be invalid.

For scientists, the peer-review process is a means by which scientific knowledge is scrutinized and determined to be meaningful and valid. Before publishing their theories, Brønsted and Lowry would have had their work examined by groups of scientists. By following the peer-review process, the Brønsted-Lowry theory was quickly accepted into the scientific body of knowledge.

Writing Brønsted-Lowry Acid-Base Reactions

Unlike the theories you have seen to this point, the Brønsted-Lowry theory attempts to describe the action of acids and bases during a chemical reaction. According to this theory, a hydrogen ion is transferred from an **acid** (the donor) to a **base** (the acceptor) during acid-base reactions. The Brønsted-Lowry theory often refers to the hydrogen ion as a proton. The products of an acid-base reaction are a **conjugate acid** and a **conjugate base**.

According to the Brønsted-Lowry Acid-Base Reactions, an acid-base reaction involves the transfer of a hydrogen ion from an acid to a base. The loss or donation of a hydrogen ion by an acid converts it into a conjugate base—another form of the substance. The conjugate base form of the substance can be recognized by the loss of a hydrogen ion in its chemical formula. The gain or acceptance of a hydrogen ion by a base converts it into a conjugate acid—its alternate form that contains the transferred hydrogen ion. The chemical formulas for many acids and bases, including their conjugate forms, are shown on the "Table of Acids and Bases." This table also appears in the Science Data Booklet, called "Relative Strengths of Selected Acids and Bases for 0.10 mol/L Solution at 25°C."

TABLE OF ACIDS AND BASES

Acid Name	Acid Formula	Conjugate Base Formula
hydrochloric acid	$HCl(aq)$	$Cl^-(aq)$
sulfuric acid	$H_2SO_4(aq)$	$HSO_4^-(aq)$
nitric acid	$HNO_3(aq)$	$NO_3^-(aq)$
hydronium ion	$H_3O^+(aq)$	$H_2O(l)$
oxalic acid	$HOOCCOOH(aq)$	$HOOCCOO^-(aq)$
sulfurous acid	$H_2SO_3(aq)$	$HSO_3^-(aq)$
hydrogen sulfate ion	$HSO_4^-(aq)$	$SO_4^{2-}(aq)$
phosphoric acid	$H_3PO_4(aq)$	$H_2PO_4^-(aq)$
orange IV	$HOr(aq)$	$Or^-(aq)$
nitrous acid	$HNO_2(aq)$	$NO_2^-(aq)$
hydrofluoric acid	$HF(aq)$	$F^-(aq)$
methanoic acid	$HCOOH(aq)$	$HCOO^-(aq)$
methyl orange	$HMo(aq)$	$Mo^-(aq)$
benzoic acid	$C_6H_5COOH(aq)$	$C_6H_5COO^-(aq)$
ethanoic (acetic) acid	$CH_3COOH(aq)$	$CH_3COO^-(aq)$
carbonic acid, $CO_2(g) + H_2O(l)$	$H_2CO_3(aq)$	$HCO_3^-(aq)$
bromothymol blue	$HBb(aq)$	$Bb^-(aq)$
hydrosulfuric acid	$H_2S(aq)$	$HS^-(aq)$
phenolphthalein	$HPh(aq)$	$Ph^-(aq)$
boric acid	$H_3BO_3(aq)$	$H_2BO_3^-(aq)$
ammonium ion	$NH_4^+(aq)$	$NH_3(aq)$
hydrogen carbonate ion	$HCO_3^-(aq)$	$CO_3^{2-}(aq)$
indigo carmine	$HIc(aq)$	$Ic^-(aq)$
water (55.5 mol/L)	$H_2O(l)$	$OH^-(aq)$

acid: the substance that donates or loses a hydrogen ion to another substance during a chemical reaction

base: the substance that accepts or gains a hydrogen ion from another substance during a chemical reaction

conjugate acid: an acid formed in an acid-base reaction when a base accepts a hydrogen ion (or proton)

conjugate base: a base formed in an acid-base reaction when an acid donates a hydrogen ion (or proton)

The information in the "Table of Acids and Bases" can be used to identify the reactants and predict the products of an acid-base reaction between certain substances.

Example Problem 1.2

Sour gas contains hydrogen sulfide, $H_2S(g)$. Hydrogen sulfide can dissolve and react with water in the atmosphere. Write the chemical equation of the reaction between aqueous hydrogen sulfide and water.

Solution

step 1: Locate $H_2S(aq)$ and $H_2O(l)$ on the "Table of Acids and Bases."

TABLE OF ACIDS AND BASES

Acid Name	Acid Formula	Conjugate Base Formula
hydrochloric acid	$HCl(aq)$	$Cl^-(aq)$
sulfuric acid	$H_2SO_4(aq)$	$HSO_4^-(aq)$
nitric acid	$HNO_3(aq)$	$NO_3^-(aq)$
hydronium ion	$H_3O^+(aq)$	$H_2O(l)$
⋮	⋮	⋮
bromothymol blue	$HBb(aq)$	$Bb^-(aq)$
hydrosulfuric acid	$H_2S(aq)$	$HS^-(aq)$
phenolphthalein	$HPh(aq)$	$Ph^-(aq)$
⋮	⋮	⋮
hydrogen carbonate ion	$HCO_3^-(aq)$	$CO_3^{2-}(aq)$
indigo carmine	$HIc(aq)$	$Ic^-(aq)$
water (55.5 mol/L)	$H_2O(l)$	$OH^-(aq)$

step 2: Identify the acid and the base in the reaction. Recall that the stronger acids appear higher in the Acid Formula column and the stronger bases appear lower in the Conjugate Base Formula column.

The acid is $H_2S(aq)$ because it appears higher in the column than $H_2O(l)$. The base is $H_2O(l)$.

step 3: Write the reactants side of the chemical equation.

$$H_2S(aq) + H_2O(l) \rightarrow$$

step 4: Identify the conjugate forms of the acid and the base.

TABLE OF ACIDS AND BASES

Acid Name	Acid Formula	Conjugate Base Formula
⋮	⋮	⋮
nitric acid	$HNO_3(aq)$	$NO_3^-(aq)$
hydronium ion	$H_3O^+(aq)$	$H_2O(l)$
⋮	⋮	⋮
bromothymol blue	$HBb(aq)$	$Bb^-(aq)$
hydrosulfuric acid	$H_2S(aq)$	$HS^-(aq)$
⋮	⋮	⋮

step 5: Write the conjugate forms on the products side of the chemical equation.

acid loses a hydrogen ion to form the conjugate base

$$H_2S(aq) + H_2O(l) \rightarrow HS^-(aq) + H_3O^+(aq)$$

base gains a hydrogen ion to form the conjugate acid

Therefore, the chemical equation for the reaction of aqueous hydrogen sulfide and water is

$$H_2S(aq) + H_2O(l) \rightarrow HS^-(aq) + H_3O^+(aq)$$

The Brønsted-Lowry theory gained acceptance within the scientific community because it was able to describe a mechanism for the reaction between acids and bases in aqueous solutions. The theory also explained the production of a hydronium ion by acids when dissolved in water.

Practice

14. Write the chemical equation for the following reactions. Label the acid, the base, the conjugate acid, and the conjugate base in each reaction.

a. Dissolved nitric acid, $HNO_3(aq)$, reacts with water, $H_2O(l)$.

b. Carbonic acid in rainwater reacts with water.

Example Problem 1.3

Hydrofluoric acid, HF(aq), used to remove oxide coatings from metals prior to electroplating, can be neutralized by a reaction with the hydroxide ion, $OH^-(aq)$, of aqueous sodium hydroxide. Write the chemical equation for this neutralization reaction.

Solution

step 1: Locate HF(aq) and $OH^-(aq)$ on the "Table of Acids and Bases."

TABLE OF ACIDS AND BASES

Acid Name	Acid Formula	Conjugate Base Formula
hydrochloric acid	$HCl(aq)$	$Cl^-(aq)$
sulfuric acid	$H_2SO_4(aq)$	$HSO_4^-(aq)$
⋮	⋮	⋮
nitrous acid	$HNO_2(aq)$	$NO_2^-(aq)$
hydrofluoric acid	$HF(aq)$	$F^-(aq)$
methanoic acid	$HCOOH(aq)$	$HCOO^-(aq)$
⋮	⋮	⋮
hydrogen carbonate ion	$HCO_3^-(aq)$	$CO_3^{2-}(aq)$
indigo carmine	$HIc(aq)$	$Ic^-(aq)$
water (55.5 mol/L)	$H_2O(l)$	$OH^-(aq)$

step 2: Identify the acid and the base in the reaction.

The acid is HF(aq), and the base is $OH^-(aq)$.

step 3: Write the reactants side of the chemical equation.

$HF(aq) + OH^-(aq) \rightarrow$

step 4: Identify the conjugate forms of the acid and the base.

acid: $HF(aq) \rightarrow F^-(aq)$

base: $OH^-(aq) \rightarrow H_2O(l)$

step 5: Write the conjugate forms on the products side of the chemical equation.

acid loses a hydrogen ion to form the conjugate base

$$HF(aq) + OH^-(aq) \rightarrow F^-(aq) + H_2O(l)$$

base gains a hydrogen ion to form the conjugate acid

Therefore, the chemical equation for this neutralization reaction is

$$HF(aq) + OH^-(aq) \rightarrow F^-(aq) + H_2O(l)$$

Practice

15. Oxalic acid, HOOCCOOH(aq), is often used in industry to clean and sterilize containers. Write the chemical equation for the reaction of oxalic acid, HOOCCOOH(aq), and the hydroxide ion, $OH^-(aq)$. Label the acid, the base, the conjugate acid, and the conjugate base.

Arrhenius's theory was not able to explain the basic properties of solutions like sodium carbonate, $Na_2CO_3(aq)$. Solutions like these are not composed of hydroxide ions; however, they can produce hydroxide ions due to a reaction with water.

Example Problem 1.4

In the "Testing Aqueous Solutions" investigation, aqueous sodium carbonate, $Na_2CO_3(aq)$, turned red litmus paper blue, indicating a basic solution. Write the chemical equation for the reaction between dissociated carbonate ions, $CO_3^{2-}(aq)$, and water.

Solution

Locate $CO_3^{2-}(aq)$ and $H_2O(l)$ on the "Table of Acids and Bases," and identify the acid and the base in the reaction.

The acid is $H_2O(l)$, and the base is $CO_3^{2-}(aq)$.

Next, write the reactants side of the chemical equation.

$$H_2O(l) + CO_3^{2-}(aq) \rightarrow$$

Now, identify the conjugate forms of the acid and the base. Write these on the products side of the chemical equation.

acid loses a hydrogen ion to form the conjugate base

$$H_2O(l) + CO_3^{2-}(aq) \rightarrow OH^-(aq) + HCO_3^-(l)$$

base gains a hydrogen ion to form the conjugate acid

Therefore, the chemical equation for the reaction is

$$H_2O(l) + CO_3^{2-}(aq) \rightarrow OH^-(aq) + HCO_3^-(aq)$$

Note: The product $OH^-(aq)$ is responsible for the basic properties of the solution.

Example Problem 1.4 shows that another aspect of the Brønsted-Lowry theory is its ability to describe the behaviour of water—either donating or accepting a hydrogen ion—and to explain its important role in acid-base reactions.

Practice

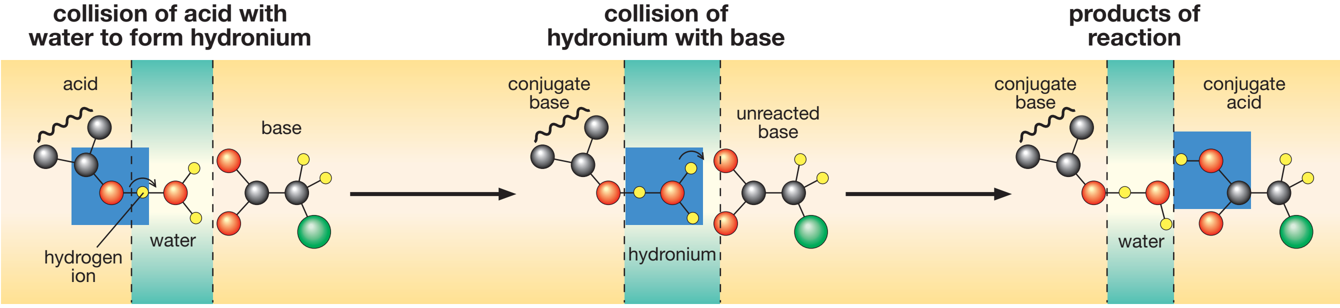
16. Write the chemical equation for each reaction given. Label the acid, the base, the conjugate acid, and the conjugate base in each reaction.

- **a.** Sulfuric acid, $H_2SO_4(aq)$, spilled during a lab procedure, reacts with the hydrogen carbonate ion, $HCO_3^-(aq)$, present within an acid spill kit.
- **b.** During the production of fertilizer, aqueous ammonia, $NH_3(aq)$, reacts with phosphoric acid, $H_3PO_4(aq)$.

17. Earlier, you determined that one of the empirical properties of acids and bases is that they act to neutralize each other. Using the Brønsted-Lowry theory, concisely explain how the neutralization of an acid or base occurs during an acid-base reaction.

Proton Hopping

Recent experiments to investigate the mechanism of hydrogen-ion transfer between acids and bases in aqueous solutions appear to have confirmed the description of the behaviour of acids and bases provided by the Brønsted-Lowry theory. Using lasers, a group of scientists were able to capture a series of images, like snapshots, of the motion of a chemical reaction between an acid and a base. The scientists expected to see the transfer of the hydrogen ion between the acid and the base present in the system, as would be described by a chemical equation similar to those you have written thus far. What the snapshots showed was unexpected: the acid and the base did not appear to come in contact with each other. The snapshots showed water molecules being converted into hydronium ions when they collided with the acid, followed by a collision between the hydronium and the base. This reaction resulted in the loss of a hydrogen ion to the base. The experiment demonstrated that water molecules in the system underwent many acid-base reactions, acting like a shuttle to transfer hydrogen ions between the acid and the base.



Although this research won't affect the way you write a reaction between an acid and a base, it provides a great deal of insight into the mechanism involved during a reaction. The research further supports the use of the Brønsted-Lowry theory when you write, explain, or predict reactions between acids and bases within an aqueous chemical system.

Emissions Can React

In Lesson 1.1 you saw how products from combustion reactions had an effect on bromothymol blue (an acid-base indicator). When the gases within exhaled air or gases from the combustion of coal were bubbled through the water containing bromothymol blue, a colour change occurred that indicated a change to an acidic system. How can bubbling gases through water result in a change to the acidity of the water?

PRODUCTS OF COMBUSTION REACTIONS AND THEIR SOURCES

Product of Combustion	Source
$CO(g)$, $CO_2(g)$	• carbon present in hydrocarbon fuels
$SO_2(g)$, $SO_3(g)$	• sulfur present in fuels • combustion of $H_2S(g)$, a component of sour gas
$NO(g)$, $NO_2(g)$	• air from the atmosphere that contains nitrogen

To summarize, products of combustion reactions (the substances shown in the table) are released into the atmosphere as emissions. Also, water is present in the atmosphere and on Earth's surface, and emissions are removed from the atmosphere in the form of wet or dry deposition. Even as dry deposition, these substances eventually come into contact with water. As you have seen, water can react with many substances.

Figure B1.18: Condensation that forms on the outside of a drinking glass was once water vapour present in the atmosphere. The presence of water and other substances makes the atmosphere a chemical system—a site for many chemical changes.

Most of the oxides of carbon, sulfur, and nitrogen shown in the "Products of Combustion Reactions and Their Sources" table can react with water. Refer to Figure B1.19.

Reactions of Certain Oxides with Water

Oxides of Carbon

carbon dioxide: $CO_2(g) + H_2O(l) \rightarrow H_2CO_3(aq)$

Oxides of Sulfur

sulfur dioxide: $SO_2(g) + H_2O(l) \rightarrow H_2SO_3(aq)$

sulfur trioxide: Sulfur dioxide, produced during combustion reactions, can convert into sulfur trioxide by reacting with oxygen present in the atmosphere. The sulfur trioxide then reacts with water to produce an acidic substance.

$$SO_2(g) + O_2(g) \rightarrow SO_3(g)$$

$$SO_3(g) + H_2O(l) \rightarrow H_2SO_4(aq)$$

Note: The term SO_x is sometimes used to refer to the presence of both $SO_2(g)$ and $SO_3(g)$ in the atmosphere.

Oxides of Nitrogen

nitrogen monoxide: Nitrogen monoxide, produced by combustion reactions, can convert into nitrogen dioxide by reacting with oxygen present in the atmosphere.

$$NO(g) + O_2(g) \rightarrow NO_2(g)$$

It is the $NO_2(g)$ that reacts with water to produce acidic substances.

nitrogen dioxide: $NO_2(g) + H_2O(l) \rightarrow HNO_2(aq) + HNO_3(aq)$

$$NO_2(g) + H_2O(l) + O_2(g) \rightarrow HNO_3(aq)$$

Note: The term NO_x refers to the presence of both NO(g) and $NO_2(g)$ in the atmosphere.

Figure B1.19

Practice

18. Copy each reaction equation listed in Figure B1.19. Balance each equation.

19. Use the "Table of Acids and Bases" on page 173 to identify the names of the hydrogen-containing products in the reaction equations. Write the name for each of the products beside its chemical formula shown in the equation.

20. The products of the reactions in Figure B1.19 often remain dissolved in water that falls toward Earth as precipitation or exists in lakes and other bodies of water. Predict the effect that the products of the reactions shown in Figure B1.19 will have on the water it dissolves in. If possible, use chemical equations to support your prediction.

anthropogenic: coming from human activity

acid deposition: airborne particles containing acids or acid-forming substances contained within precipitation (wet deposition) or that absorb directly into parts of Earth's surface (dry deposition)

Acid Deposition

The chemical reactions you have studied thus far describe the origin and consequences of substances released as emissions from human activity. **Anthropogenic** emissions of carbon dioxide, $CO_2(g)$, sulfur oxides, $SO_2(g)$ and $SO_3(g)$, NO_x or nitrous oxides, NO(g) and $NO_2(g)$, originate from human-made processes that involve combustion, such as energy production and transportation. The chemical equations you have written describe how these emissions are able to react with water and form **acid deposition**. Areas exposed to wet or dry acid deposition can experience a number of effects, some of which you may be able to predict based on what you have already learned about acids from previous investigations.

Figure B1.20: Water from melting snow can contain acids that were originally deposited in either wet or dry form.

The term **acid rain** describes the excessive amount of acidity within precipitation. It surprises many people to learn that rainfall is naturally acidic. Natural processes, like cellular respiration, produce substances that can form acids. Carbon dioxide, $CO_2(g)$, for example, can react with water to form carbonic acid, $H_2CO_3(aq)$, which results in the production of hydronium ions, $H_3O^+(aq)$. The oxides of carbon, nitrogen, and sulfur are present in Earth's atmosphere as a result of natural processes and anthropogenic sources. Other natural processes that release emissions include burning biomass (forest fires), lightning, the erosion of carbonate-based rock formations, the release of volcanic gases, and the action of bacteria in marine or terrestrial habitats.

acid rain: any form of precipitation (wet deposition) containing an excess of dissolved acids; wet deposition with a pH of 5.6 or less

The amount of acid present within a solution can be measured. The method used can determine whether precipitation contains higher amounts of acids than would be expected. In many areas of the world, acid deposition is largely due to the emissions from human activities. The amount of acid deposition in a region can vary over time, even over seasons. Can you predict the effect that melting snow has on the amount of acid present in streams and bodies of water in Alberta?

Practice

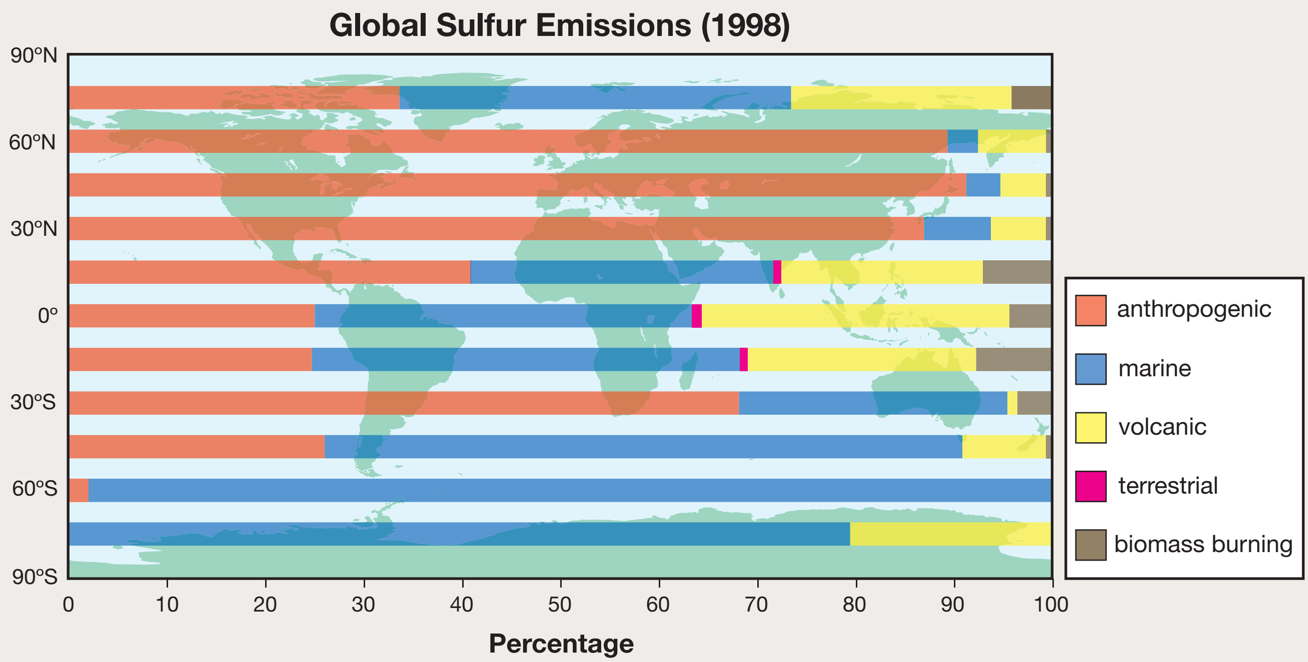
21. Analyze the graph "Global Sulfur Emissions (1998)."

a. Prepare a table that lists the following values by latitude:

- percentage of emissions from anthropogenic sources
- percentage of emissions from natural sources

b. Calculate a ratio of anthropogenic sources to natural sources for each latitude zone shown and for the world.

c. Identify areas where you suspect higher levels of acid deposition may occur.



Measuring Acids

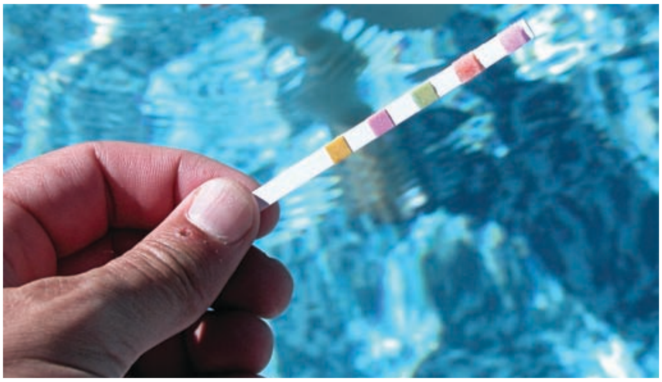


Figure B1.21: The level of acidity in a swimming pool is checked by measuring the water's pH.

During the hot summer days, the place to be is at the local swimming pool. But before anyone can go in the water, lifeguards must take measurements of the water's acidity to ensure that it will not cause any harm. Acids are corrosive. Acidic solutions react with metals and often have warning labels to remind you of their ability to react with your skin.

Figure B1.22: Labelling products using WHMIS symbols or HHPS (Household Hazardous Products Symbols) alert users of the possible risks and appropriate safety precautions.

pH

pH: a value that represents the concentration of dissolved hydronium ions, $H_3O^+(aq)$, within a solution

One way to measure the amount of acid in a solution—expressed as a concentration of hydronium ions, $H_3O^+(aq)$—is by measuring the solution's pH. Concentrated acidic solutions contain a larger number of moles of hydronium ions within each millilitre or litre of solution than dilute acidic solutions. You may have noticed that some of the labels on cleaning products containing acids or bases in your home indicate that they are "concentrated." If the same amount of concentrated cleaner and dilute cleaner were tested, the concentrated solution should remove a stain more quickly than the dilute solution because of the higher number of particles available to react with the molecules in the affected area. Recall that the presence of hydronium ions, $H_3O^+(aq)$, gives a solution its acidic properties. The concentration of hydronium ions within an acidic solution influences other aspects of the reaction involving acids, including

- how quickly the solution will begin to react
- how much change the acid may cause
- the amount of base required to neutralize the acid
- the amount of base or metal it will react with

In Alberta, the pH of rainfall is routinely measured at a number of locations throughout the province to provide information about acid deposition.

Calculating pH and the pH Scale

The pH scale was developed in 1909 by a Danish scientist named Soren Sörenson. He developed the scale as a means to better communicate the acidity of a solution. Scientists at that time observed that the level of acidity of a solution did not always correspond to the concentration of the acid dissolved in the solution. Not all acids react completely with water, thus producing solutions with lower concentrations of hydronium ions. Sörenson's system was designed to measure the concentration of hydronium ions present in dilute solutions, thereby providing a better description of a solution's level of acidity. A solution with a pH of 7 is considered to be neutral; a solution with a pH less than 7 is considered to be acidic; and a solution with a pH greater than 7 is considered to be basic. Figure B1.23 shows the pH scale along with examples of common substances.

pH = 0	battery acid
pH = 1	stomach acid
pH = 2	lemon juice
pH = 3	vinegar, orange juice, cola
pH = 4	tomato juice, acid rain
pH = 5	coffee (black), rain
pH = 6	urine, saliva (healthy), cow's milk
pH = 7	distilled water, human blood
pH = 8	sea water
pH = 9	baking soda
pH = 10	milk of magnesia, detergent
pH = 11	ammonia solution, household cleaners
pH = 12	hand soap
pH = 13	bleach, oven cleaner, household lye
pH = 14	liquid drain cleaner

Figure B1.23: pH scale

The exponent of the hydronium-ion concentration, when expressed in scientific notation, can be used to approximate a solution's pH. For example, a pH of 6.0 corresponds with a hydronium-ion concentration of 1×10^{-6} mol/L, which can also be expressed as 0.000 001 mol/L. A solution with a pH of 6.0 has a larger concentration of hydronium ions than a solution with a pH of 7.0, which corresponds to a hydronium-ion concentration of 1×10^{-7} mol/L or 0.000 000 1 mol/L. By dividing the hydronium-ion concentration of a solution with a pH of 6.0 by the hydronium-ion concentration of a solution with a pH of 7.0, you will see that the solution with the pH of 6.0 has ten times more hydronium ions than the solution with the pH of 7.0.

$$\begin{aligned}\frac{\text{solution with pH 6.0}}{\text{solution with pH 7.0}} &= \frac{H_3O^+ \text{concentration}}{H_3O^+ \text{concentration}}\\ &= \frac{1\times10^{-6}\ \text{mol/L}}{1\times10^{-7}\ \text{mol/L}}\\ &= 10\end{aligned}$$

Each whole-number division on the pH scale represents a ten-fold difference in the concentration of hydronium ions from the value above or below it. As you move down the scale toward higher pH values, the hydronium-ion concentration decreases, and vice versa as you move up the scale toward lower pH values. A change of two pH steps on the scale represents two ten-fold changes, or a 100-fold change, to the hydronium-ion concentration. If you compare the hydronium-ion concentration of a solution with a pH of 4 with that of a solution with a pH of 9, there is a 100 000-fold $\left(10^5\right)$ difference in hydronium-ion concentration.

The pH scale was developed using dilute solutions of acids and bases. The concentration of hydronium ions within a dilute solution is small and is often expressed in scientific notation as an exponent to the base 10. Because logarithms calculate exponents, you can use the logarithm function on your calculator. Use the following equation to calculate pH.

$$\text{pH} = -\log_{10}\left[H_3O^+(aq)\right]$$

The pH of concentrated acid solutions can be below the range of the pH scale. A pH of 0 corresponds to a solution containing a hydrogen-ion concentration of 1.00 mol/L (1×10^0 mol/L). Concentrated stock acid solutions used in laboratories can have negative pH values.

Example Problem 1.5

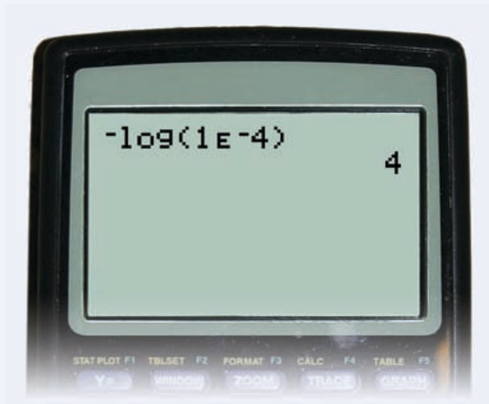
Determine the pH of a sample of rainwater that has a hydronium-ion concentration, $\left[H_3O^+(aq)\right]$, of 1.00×10^{-4} mol/L.

Solution

$$\begin{aligned}\text{pH} &= -\log_{10}\left[H_3O^+(aq)\right]\\ &= -\log_{10}\left(1.00 \times 10^{-4}\ \text{mol/L}\right) \leftarrow \text{Substitute the hydronium-ion concentration.}\\ &= 4\end{aligned}$$

Keystrokes for Graphing Calculator

(–) LOG 1 2nd [EE] (–) 4) ENTER



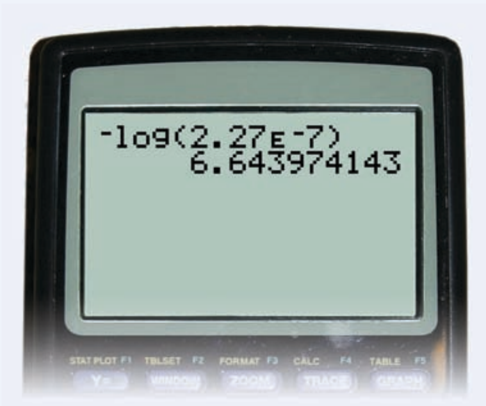
The pH of the rainwater is 4.000, expressed to the appropriate number of significant digits.

Example Problem 1.6

A sample of lake water has a hydronium-ion concentration of 2.27×10^{-7} mol/L. Determine the pH of the lake water.

Solution

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+(\text{aq})]$$
$$= -\log_{10}(2.27 \times 10^{-7}\ \text{mol/L})$$
$$= 6.644 \quad \leftarrow \text{3 significant digits}$$



The lake water has a pH of 6.644.

Significant Digits and pH Calculations

Using logarithms to represent calculations is convenient, but care must be taken to show the appropriate number of significant digits. The number to the left of the decimal point in a pH value represents the order of magnitude of the hydronium-ion concentration, reflected by the exponent of the base 10 when that concentration is written in scientific notation. The exponent provides information only about how large or small the number is, not how accurately it was measured. Accuracy of measurement is determined by examining the other numbers before and after the decimal point. Significant digits in a pH value are written using the appropriate number of digits after the decimal point.

Significant Digits and pH Values

$[\text{H}_3\text{O}^+(\text{aq})]$ concentration = 0.000 010 mol/L

$$= \underbrace{1.0}_{\text{2 significant digits}} \times 10^{-5}\ \text{mol/L}$$

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+(\text{aq})]$$
$$= -\log_{10}(1.0 \times 10^{-5}\ \text{mol/L})$$
$$= 5.\underbrace{00}_{\text{2 significant digits}}$$

Calculating Hydronium Ions

An algebraic relationship in mathematics, such as a formula, can be rearranged to solve for any variable in the formula. Therefore, the formula to calculate pH can be rearranged to solve for the hydronium-ion concentration when a solution's pH value is known.

$$\text{pH} = -\log_{10}[\text{H}_3\text{O}^+(\text{aq})]$$
$$[\text{H}_3\text{O}^+(\text{aq})] = 10^{-\text{pH}}$$

Example Problem 1.7

Calculate the hydronium-ion concentration, $[\text{H}_3\text{O}^+(\text{aq})]$ in a shampoo with a pH of 5.72.

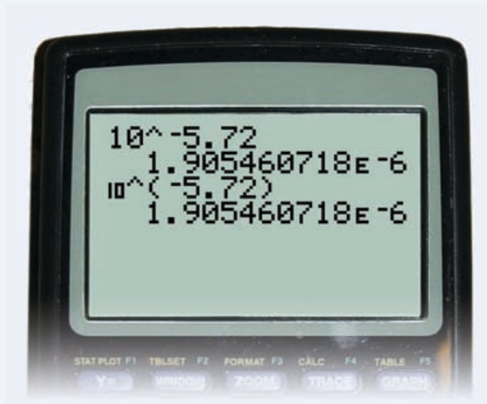
Solution

$$[\text{H}_3\text{O}^+(\text{aq})] = 10^{-\text{pH}}$$
$$= 10^{-5.72} \quad \leftarrow \text{2 significant digits in pH values}$$
$$= 1.905\,460\,718 \times 10^{-6}\ \text{mol/L}$$
$$= 1.9 \times 10^{-6}\ \text{mol/L} \quad \leftarrow \text{2 signficant digits}$$

Keystrokes for Graphing Calculator

Method 1: 10 [^] [(–)] 5.72 [ENTER]

Method 2: [2nd] [10^x] [(–)] 5.72 [)] [ENTER]



The shampoo has a hydronium-ion concentration of 1.9×10^{-6} mol/L.

The formula for pH and the formula for hydronium-ion concentration appear in the Science Data Booklet.

Practice

22. For each hydronium-ion concentration, calculate the pH. Classify each solution as being acidic, basic, or neutral.
    - **a.** 0.001 00 mol/L
    - **b.** 2.00×10^{-4} mol/L
    - **c.** 1.5×10^{-6} mol/L
    - **d.** 1.35×10^{-8} mol/L
    - **e.** 1.54×10^{-12} mol/L
23. For each pH value, calculate the hydronium-ion concentration.
    - **a.** 7.00
    - **b.** 2.98
    - **c.** 8.912
    - **d.** 13.1
24. Obtain the handout "pH and Hydronium-Ion Concentration" from the Science 30 Textbook CD. Complete the table by calculating values for the $[H_3O^+(aq)]$ column and the Relative Change in Hydronium-Ion Concentration with pH Value Below column. State a general trend in the last column regarding the change to hydronium-ion concentration along the pH scale.

Investigation

Measuring pH Using Indicators

Purpose

You will demonstrate the use of indicators as a means to determine the pH of a solution.

Science Skills

✓ Performing and Recording
✓ Analyzing and Interpreting

Materials

- 2 copies of the handout "Determining pH Using Indicators" from the Science 30 Textbook CD
- 1 letter-size overhead transparency sheet
- masking tape
- dropper bottles (or eyedroppers)
- solutions with pHs of 1, 3, 5, 7, 9, 11, and 13
- pH indicators
  - alizarin yellow R
  - thymol blue
  - bromothymol blue
  - bromocresol green
  - methyl orange
- water containing juice extracted from red cabbage when boiled
- unknown solutions A, B, and C

CAUTION!

Use gloves, safety glasses, and a lab apron for this activity.

Procedure

step 1: Place one of the handouts on the surface of your work area. Cover the handout with the transparency sheet. Use masking tape to ensure the sheet lays flat and remains attached to the surface of your work area throughout the experiment.

step 2: Place one drop of the solution labelled "pH 1" into each circle in its designated column on the transparency, which is overtop the handout.

step 3: Repeat step 2 with the other pH solutions and the unknown solutions.

step 4: Carefully add a drop of alizarin yellow R indicator to the circles in the first row of the handout. When adding the indicator, ensure that the end of the bottle does not touch the drop of solution already in the circle.

step 5: Repeat step 4 using the other indicators listed in the handout.

step 6: Record the colour of the resulting mixture within each circle in a data table.

step 7: Use paper towel to absorb most of the solutions on the transparency; then rinse the transparency in the sink.

step 8: Return all apparatus to their proper location in the lab.

Observations

1. Show your results of this investigation.

Analysis

2. Use the data to estimate the pH of solutions A, B, and C. Explain how you arrived at your estimation. State a reason why using indicators results in only an estimation of the pH of the three solutions.

Use of Natural Indicators by First Nations

The name for the Blackfoot First Nation is reported to have originated from the black moccasins worn by their members when first encountered by European settlers. Research has shown that the Blackfoot used over 150 different plant species to support their traditional lifestyle. The skunkbush, as well as other prairie plants, were sources of coloured molecules that could be used to dye leather, cloth, and even the porcupine quills for ceremonial dress. The Blackfoot, as well as other First Nations, used naturally occurring acids to adjust the colour of the dyes made from the extracts of berries, leaves, or bark. They used ash from fire pits because ash combined with water forms a basic solution. Changing colour as a response to differing pH is a property of many natural substances, making them useful **indicators** of changes in the pH of a system that are often the result of an acid-base reaction.

Figure B1.24: A black dye can be prepared from the leaves of the skunkbush plant.

indicator: a substance that changes colour in response to the change in pH of a system

Using Indicators to Estimate pH

The colours shown by many indicators at various pH values are summarized in the "Acid-Base Indicators" table. The information in this table can be used to interpret and estimate the pH of a solution. This table also appears in the Science Data Booklet.

ACID-BASE INDICATORS

Indicator	Abbreviation (acid / conjugate base)	pH Range	Colour Change as pH Increases
methyl violet	$HMv(aq)/Mv^-(aq)$	0.0 – 1.6	yellow to blue
thymol blue	$H_2Tb(aq)/HTb^-(aq)$	1.2 – 2.8	red to yellow
thymol blue	$HTb^-(aq)/Tb^-(aq)$	8.0 – 9.6	yellow to blue
orange IV	$HOr(aq)/Or^-(aq)$	1.4 – 2.8	red to yellow
methyl orange	$HMo(aq)/Mo^-(aq)$	3.2 – 4.4	red to yellow
bromocresol green	$HBg(aq)/Bg^-(aq)$	3.8 – 5.4	yellow to blue
litmus	$HLt(aq)/Lt^-(aq)$	4.5 – 8.3	red to blue
methyl red	$HMr(aq)/Mr^-(aq)$	4.8 – 6.0	red to yellow
chlorophenol red	$HCh(aq)/Ch^-(aq)$	5.2 – 6.8	yellow to red
bromothymol blue	$HBb(aq)/Bb^-(aq)$	6.0 – 7.6	yellow to blue
phenol red	$HPr(aq)/Pr^-(aq)$	6.6 – 8.0	yellow to red
phenolphthalein	$HPh(aq)/Ph^-(aq)$	8.2 – 10.0	colourless to pink
thymolphthalein	$HTh(aq)/Th^-(aq)$	9.4 – 10.6	colourless to blue
alizarin yellow R	$HAy(aq)/Ay^-(aq)$	10.1 – 12.0	yellow to red
indigo carmine	$HIc(aq)/Ic^-(aq)$	11.4 – 13.0	blue to yellow
1,3,5–trinitrobenzene	$HNb(aq)/Nb(aq)$	12.0 – 14.0	colourless to orange

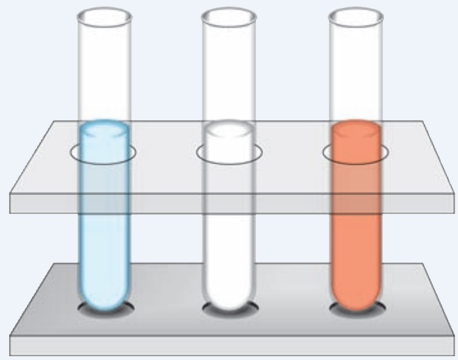
At the beginning of this unit, you observed a colour change for bromothymol blue, one of the pH indicators listed in the "Acid-Base Indicators" table. As you may recall, the colour of the indicator in the water solution within the flask was either blue or green. As gases from exhaled air and from the products of the combustion of coal were added to the respective flasks, the colour of the indicator within the water turned yellow. As can be interpreted from the "Acid-Base Indicators" table, the yellow colour observed for the flask that contained bromothymol blue indicates that the pH of the solution must have been below 6.0. Earlier in the demonstration, the blue colour indicated that the pH of the water in the flask was above 7.6. (A green colour would have indicated that the pH of the water in the flask was between 6.0 and 7.6.)

In the "Measuring pH Using Indicators" investigation, observations from more than one indicator were needed to estimate the pH of a solution to a reasonable degree of accuracy. When an observation can only estimate that a solution's pH is below 6, additional indicator data must be used to narrow the range for a more accurate estimate. Example Problem 1.8 demonstrates how data from several indicators can be used to determine the pH of a solution.

Example Problem 1.8

A solution is tested with three indicators. Here are the results.

Indicator	Colour
bromothymol blue	blue
phenolphthalein	colourless
phenol red	red



Estimate the pH of this solution.

Solution

step 1: Use the “Acid-Base Indicators” table to determine the pH range.

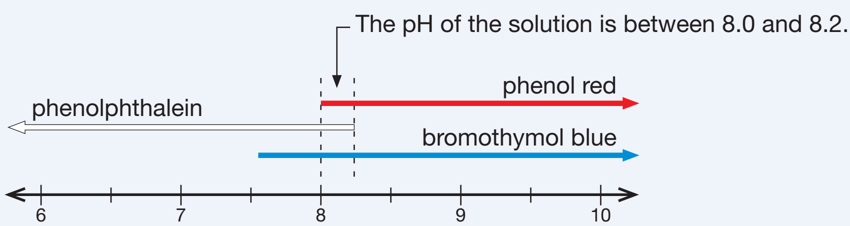
ACID-BASE INDICATORS

Indicator	Abbreviation (acid/conjugate base)	pH Range	Colour Change as pH Increases
methyl violet	$HMv(aq)/Mv^-(aq)$	0.0 – 1.6	yellow to blue
thymol blue	$H_2Tb(aq)/HTb^-(aq)$	1.2 – 2.8	red to yellow
⋮	⋮	⋮	⋮
bromothymol blue	$HBb(aq)/Bb^-(aq)$	6.0 – 7.6	yellow to blue
phenol red	$HPr(aq)/Pr^-(aq)$	6.6 – 8.0	yellow to red
phenolphthalein	$HPh(aq)/Ph^-(aq)$	8.2 – 10.0	colourless to pink
⋮	⋮	⋮	⋮

According to the table, the three indicators show the following:

- **bromothymol blue:** blue = above pH 7.6
- **phenolphthalein:** colourless = below pH 8.2
- **phenol red:** red = above pH 8.0

step 2: Use a number line to narrow the range of pH values.



The estimated pH of the solution is between 8.0 and 8.2.

A pH meter is used when more accurate measurements of the pH of solutions are required. A pH meter contains a probe that detects the concentration of hydronium ions in a solution. The sensitivity of the probe enables pH meters to measure values to the hundredth or thousandth of a pH unit.

Figure B1.25: pH meters are used in the laboratory and in field work to accurately measure the pH of rainwater and other solutions.

Practice

25. Describe what an indicator is and how it can determine the pH of a solution.

26. "Methyl orange is always red in an acidic solution and yellow in a basic solution." Explain whether this statement is correct or incorrect.

27. A solution is tested with three indicators. Here are the results.

Indicator	Colour
litmus	red
thymol blue	orange
orange IV	red

a. Estimate the pH of the solution.

b. Estimate the hydronium-ion concentration of the solution.

28. A few drops of bromocresol green and a few drops of thymol blue are added to the same solution with a pH of 5.6. What colour do you expect the solution to be? Concisely explain your answer.

Sources of Acid Deposition

Earlier, equations demonstrated how carbon dioxide, sulfur oxides, and nitrous oxides reacted with water to produce hydronium ions and form acidic solutions. Carbon dioxide—although produced by many combustion processes—is not considered a major source of acid deposition. Carbon dioxide is only slightly soluble in water, which limits the extent to which it can react with water. On the other hand, the oxides of nitrogen and sulfur are considerably more soluble in water and, thus, make a greater contribution to the acidification of wet and dry deposition.

1.2 Summary

In this lesson you examined the chemistry of acids and bases. You wrote chemical equations to describe the changes that occur when acids and bases react and discovered how substances released into the atmosphere can convert into acids. You discovered that chemical reactions that occur between acids and water produce hydronium ions. You also determined that the substances released as emissions can return to Earth's surface as either wet or dry deposition and that the deposition can be acidic. You then measured the pH of a solution and studied how pH relates to the concentration of hydronium ions within a solution. In the next lesson you will cover the effects of acidic deposition on the environment.

1.2 Questions

Knowledge

1. Define the following terms.
   - **a.** acid
   - **b.** base
   - **c.** dissociation
   - **d.** hydrogen ion
   - **e.** hydronium ion
   - **f.** pH scale
   - **g.** wet deposition
   - **h.** acid deposition
   - **i.** acid rain
2. Write balanced chemical equations for the reactions between the following substances. For each equation, label the acid, the base, the conjugate acid, and the conjugate base.
   - **a.** hydronium ion, $H_3O^+(aq)$, and hydroxide ion, $OH^-(aq)$
   - **b.** ethanoic acid and ammonia
3. List similarities and differences between Arrhenius's theory and the Brønsted-Lowry theory.
4. Calculate the pH values for each hydronium-ion concentration given. Identify whether the solution is acidic, basic, or neutral.
   - **a.** 0.001 25 mol/L
   - **b.** 2.3×10^{-9} mol/L
   - **c.** 4.42×10^{-13} mol/L
   - **d.** 5.6×10^{-2} mol/L
   - **e.** 8.10×10^{-8} mol/L
5. Calculate hydronium-ion concentration for each pH value given.
   - **a.** 2.14
   - **b.** 7.1
   - **c.** 9.437
   - **d.** 11.00

Applying Concepts

6. Compare and contrast the terms *proton*, *hydrogen ion*, and *hydronium ion*.
7. Antacids are usually taken to relieve heartburn. State the type of compound an antacid needs to be in order to be effective. Calcium carbonate, $CaCO_3(s)$, and aluminium hydroxide, $Al(OH)_3(s)$, are substances used in commercially available antacids. List the empirical properties common to these two antacids. Write a balanced chemical equation that represents the reaction between each of these antacids and aqueous hydronium ions that would occur in the stomach.
8. A chemical spill releases concentrated ammonia, $NH_3(aq)$, along a dangerous-goods route. The spill has been contained. Identify the general properties of the concentrated ammonia spill. If a decision is made to treat the spill to reduce the risk to people or the environment, indicate a substance that can be used. Support your answer with a balanced chemical equation.
9. A solution is yellow with thymol blue and blue with bromocresol green. Determine the colour of the solution with the following indicators.
   - **a.** methyl violet
   - **b.** indigo carmine
   - **c.** methyl orange
   - **d.** alizarin yellow R
10. "The total amount of acid being deposited in an area is equal to the amount of wet acidic deposition deposited in the area plus the amount of dry acidic deposition deposited in the area." Use the concepts you applied in this lesson to explain whether you think this statement is correct or incorrect.
11. Identify whether each example affects the validity or reliability of scientific work.
    - **a.** repeating an experiment
    - **b.** comparing your data with the data collected by other students completing the same experiment
    - **c.** two groups of scientists arriving at the same result using different methods
12. Refer to the table you prepared in Practice Problem 11 on page 169. Complete the Definition column.